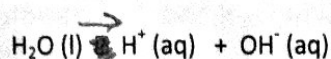


T08D04 : pH and the Equilibrium Constant for Water

Name

Part A: Essentials

Water will ionize(dissociate) into its ions according to the following equation



The equilibrium constant for water is therefore

$$K_w = [\text{H}^+][\text{OH}^-] \text{ and its value is } =$$

1×10^{-14} at 25 degrees C. This value means that whether there is an excess of H^+ , OH^- , or equal amounts, the sum of the parts is still equal to 1×10^{-14} .

In comparison the equilibrium value for HCl is exceedingly large $\text{H}^+ \gg \text{HCl}$

Part B: The Formulas and Problems

The formula for pH is :

$$-\log [\text{H}^+] = \text{pH} \quad \times \text{ where } [\text{H}^+] \text{ is the concentration of } \text{H}^+ \text{ ions in mol/L (molarity for strong acids)}$$

The additional formulas derived from the pH and K_w formulas are:

$$\begin{aligned} \text{pH} + \text{pOH} &= 14 \\ -\log [\text{OH}^-] &= \text{pOH} \\ 10^{-\text{pH}} &= [\text{H}^+] \quad 10^{-\text{pOH}} = [\text{OH}^-] \end{aligned}$$

1. Using this formula calculate the pH of the following:

- | | | | |
|---------------------|-----------------------|-------|----------------------------------|
| a. $[\text{H}^+] =$ | 1.67×10^{-4} | 3.78 | $-\log [\text{H}^+] = \text{pH}$ |
| b. $[\text{H}^+] =$ | 1.1×10^{-5} | 4.96 | |
| c. $[\text{H}^+] =$ | 1.8×10^{-11} | 10.74 | |
| d. $[\text{H}^+] =$ | 5.5×10^{-6} | 5.26 | |

2. Calculate the pH of the solutions that have the following concentrations of hydrogen or hydroxide ions:

- | | | | |
|----------------------|-----------------------|-------|----------------------------------|
| a. $[\text{H}^+] =$ | 1.67×10^{-4} | 3.78 | $-\log [\text{H}^+] = \text{pH}$ |
| b. $[\text{OH}^-] =$ | 1.1×10^{-5} | 9.04 | |
| c. $[\text{H}^+] =$ | 1.8×10^{-11} | 10.74 | |
| d. $[\text{OH}^-] =$ | 5.5×10^{-6} | 8.74 | |
- $14 - \text{pOH} = \text{pH}$

3. Calculate the hydrogen ion concentration, $[H^+]$, and the hydroxide ion concentration, $[OH^-]$ of each of the following solutions given their pH:

	pH	pOH	$[H^+]$	$[OH^-]$
a.	2.9	11.1	1.26×10^{-3}	7.94×10^{-12}
b.	5.5	8.5	3.16×10^{-6}	3.16×10^{-9}
c.	10.1	3.9	7.94×10^{-11}	1.26×10^{-4}
d.	-1.4	~14	$\sim 1 \times 10^0$	$\sim 1 \times 10^{-14}$

4. Fill in the following table:

$[H^+]$	$[OH^-]$	pH	pOH
1.26×10^{-2}	7.94×10^{-13}	1.9	12.1
6.67×10^{-11}	1.5×10^{-4}	10.18	3.82
1×10^{-14}	1.0	14	-0.60
2.3×10^{-2}	4.35×10^{-13}	1.64	12.36
3.16×10^{-6}	3.16×10^{-9}	5.5	8.5
0.000155	6.45×10^{-11}	3.81	10.19

5. Calculate the pH, pOH, $[H^+]$, and $[OH^-]$ concentration of each of the following:

a. 1.0 M solution of HNO_3 - acid

$$-\log(1.0) = pH = 0 \quad [H^+] = 1 \times 10^0$$

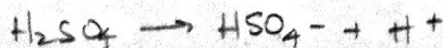
$$pOH = 14 \quad [OH^-] = 1 \times 10^{-14}$$

b. 0.65 M solution of $NaOH$ - base

$$-\log(0.65) = pOH = 0.19 \quad [OH^-] = 6.5 \times 10^{-1}$$

$$pH = 13.81 \quad [H^+] = 1.5 \times 10^{-14}$$

c. 13.45 g of H_2SO_4 placed in water to make 325 mL of solution. [Diprotic acid but only the first ionization step should be used.]



$$13.45 \text{ g } H_2SO_4 \times \frac{1 \text{ mol } H_2SO_4}{98.09 \text{ g } H_2SO_4} = 0.137 \text{ mol } H_2SO_4$$

$$\frac{0.137 \text{ mol}}{0.325 \text{ L}} = \frac{\text{mol}}{\text{L}} = \text{Molarity}$$

$$= 0.422 \text{ M}$$

$$-\log(0.442) = 0.38 = pH \quad [H^+] = 0.442$$

$$13.62 = pOH \quad [OH^-] = 2.26 \times 10^{-14}$$